National 5 Chemistry

Unit 1:

Chemical Changes & Structure

Student:

Topic 4
Chemistry of Acids & Bases

Topics

<table>
<thead>
<tr>
<th>Sections</th>
<th>Done</th>
<th>Checked</th>
</tr>
</thead>
<tbody>
<tr>
<td>4.1 Acids &amp; Bases</td>
<td></td>
<td></td>
</tr>
<tr>
<td>1. Common Acids</td>
<td></td>
<td></td>
</tr>
<tr>
<td>2. Common Bases</td>
<td></td>
<td></td>
</tr>
<tr>
<td>3. Bases &amp; Alkalis</td>
<td></td>
<td></td>
</tr>
<tr>
<td>4. Making Acids</td>
<td></td>
<td></td>
</tr>
<tr>
<td><strong>Self-Check Questions 1 - 8</strong></td>
<td>Score:</td>
<td>/</td>
</tr>
<tr>
<td>4.2 Acid &amp; Base Structures</td>
<td></td>
<td></td>
</tr>
<tr>
<td>1. Acid Molecules</td>
<td></td>
<td></td>
</tr>
<tr>
<td>2. Covalent to Ionic</td>
<td></td>
<td></td>
</tr>
<tr>
<td>3. Ammonia</td>
<td></td>
<td></td>
</tr>
<tr>
<td>4. Water Ions - Dissociation</td>
<td></td>
<td></td>
</tr>
<tr>
<td>5. pH Numbers</td>
<td></td>
<td></td>
</tr>
<tr>
<td><strong>Self-Check Questions 1 - 8</strong></td>
<td>Score:</td>
<td>/</td>
</tr>
<tr>
<td>4.3 Reactions of Acids</td>
<td></td>
<td></td>
</tr>
<tr>
<td>1. With Metals</td>
<td></td>
<td></td>
</tr>
<tr>
<td>2. With Metal Oxides</td>
<td></td>
<td></td>
</tr>
<tr>
<td>3. With Metal Carbonates</td>
<td></td>
<td></td>
</tr>
<tr>
<td>4. With Alkalis</td>
<td></td>
<td></td>
</tr>
<tr>
<td>5. Making Salts</td>
<td></td>
<td></td>
</tr>
<tr>
<td><strong>Self-Check Questions 1 - 8</strong></td>
<td>Score:</td>
<td>/</td>
</tr>
<tr>
<td>4.4 Quantitative Analysis</td>
<td></td>
<td></td>
</tr>
<tr>
<td>1. Standard Solution</td>
<td></td>
<td></td>
</tr>
<tr>
<td>2. Titrations</td>
<td></td>
<td></td>
</tr>
<tr>
<td>3. Technique</td>
<td></td>
<td></td>
</tr>
<tr>
<td>4. Results</td>
<td></td>
<td></td>
</tr>
<tr>
<td>5. Evaluation</td>
<td></td>
<td></td>
</tr>
<tr>
<td><strong>Self-Check Questions 1 - 8</strong></td>
<td>Score:</td>
<td>/</td>
</tr>
<tr>
<td>Consolidation Work</td>
<td></td>
<td></td>
</tr>
<tr>
<td><strong>End-of-Topic Assessment</strong></td>
<td>Score:</td>
<td>%</td>
</tr>
<tr>
<td><strong>Grade:</strong></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>
Knowledge Met in this Topic

Common household acids and alkalis

- **Acids**: vinegar, citrus fruits, cola drinks etc
- **Alkalis**: lime, oven cleaner, bleach, bicarbonate of soda, soap, ammonia

Oxides and hydroxides

- Oxides of **non-metals** which dissolve produce **acidic** solutions e.g. CO₂, SO₂ and NO₂.
- Non-metal oxides are the main cause of **acid rain**.
- Oxides and hydroxides of **metals** which dissolve produce **alkaline** solutions.
- All the oxides of Group 1 metals are very soluble, only some from Group 2 are soluble.

Important acids

- Most acids start off as **covalent molecules** which break apart (**dissociate**) in water to produce **hydrogen ions**, H⁺
- **hydrochloric** acid: HCl → H⁺ + Cl⁻
- **sulphuric** acid: H₂SO₄ → 2H⁺ + SO₄²⁻
- **nitric** acid: HNO₃ → H⁺ + NO₃⁻

The pH scale

- pH is a number that shows how acidic or alkaline a solution is.
- Universal indicator, pH paper or a pH meter can show the pH of a solution.
- **Acids**: pH less than 7, pH < 7
  - **Neutral**: pH equals 7, pH = 7
  - **Alkalis**: pH more than 7, pH > 7
- When acids dissolve in water they produce hydrogen ions, H⁺
- When alkalis dissolve in water they produce hydroxide ions, OH⁻
- Pure water and all neutral solutions contain a tiny but equal concentration of hydrogen and hydroxide ions, H⁺ = OH⁻
- An acid solution contains more hydrogen ions than hydroxide, H⁺ > OH⁻
- An alkali solution contains less hydrogen ions than hydroxide, H⁺ < OH⁻
- Diluting acids or alkalis will reduce the concentration of H⁺ and OH⁻, and move the pH towards 7.

Neutralisation

- **Neutralisation** is a reaction in which the pH of a solution moves towards 7.


Everyday examples of neutralisation

- Lime (calcium oxide) is used to reduce acidity in soil and water.
- Cures for acid indigestion contain neutralisers such as calcium carbonate.

Bases and alkalis

- A base is a substance that neutralises an acid.
- An alkali is a base that dissolves in water.

Salts

- Salts are ionic compounds formed in reactions between acids and neutralisers.
- A metal ion or an ammonium ion will have replaced the hydrogen in an acid.
- Hydrochloric acid (HCl) forms chlorides e.g. NaCl
- Sulphuric acid (H₂SO₄) forms sulphates e.g. CuSO₄
- Nitric acid (HNO₃) forms nitrates e.g. NH₄NO₃
- All the oxides of Group 1 metals are very soluble, only some from Group 2 are soluble.

Acid Reactions

- Acid + Alkali → Water + ‘salt’
  eg \( H^+ \) Cl\(^-\) + Na\(^+\) OH\(^-\) → H\(_2\)O + Na\(^+\) Cl\(^-\)
  removing spectator ions:- \( H^+ \) + OH\(^-\) → H\(_2\)O

- Acid + Metal → Hydrogen + ‘salt’
  eg \( (H^+) \)\(_2\)SO₄\(^-\) + Mg → H\(_2\) + Mg\(^{2+}\)SO₄\(^-\)
  removing spectator ions:- 2H\(^+\) + Mg → H\(_2\) + Mg\(^{2+}\)

- Acid + Oxide/Hydroxide → Water + ‘salt’
  eg 2H\(^+\) NO₃\(^-\) + Cu\(^{2+}\) O\(^2-\) → H\(_2\)O + Cu\(^{2+}\) (NO₃\(^-)\)\(_2\)
  removing spectator ions:- 2H\(^+\) + O\(^2-\) → H\(_2\)O

- Acid + Carbonate → Water + Carbon dioxide + ‘salt’
  eg 2H\(^+\) Cl\(^-\) + Ca\(^{2+}\) CO₃\(^-\) → H\(_2\)O + CO\(_2\) + Ca\(^{2+}\) (Cl\(^-\))\(_2\)
  removing spectator ions:- 2H\(^+\) + CO\(_3\)\(^-\) → H\(_2\)O + CO\(_2\)

- Acid rain reacts with carbonate rocks such as marble (statues) and limestone, and with metals such as iron.
- Reacting acids is a good way of making salts.
- Salts can also be made by precipitation.
**Ammonia**

- Ammonia is a colourless gas with a sharp, unpleasant (pungent) smell.
- Ammonia is a very soluble in water producing an **alkaline** solution.

\[ \text{NH}_3 + \text{H}_2\text{O} \rightarrow \text{NH}_4^+ + \text{OH}^- \]

- Ammonia is the **only** common alkaline gas.
- Ammonia can neutralise an acid and form an ammonium salt.

\[ \text{NH}_3 + \text{H}_2\text{SO}_4 \rightarrow (\text{NH}_4)_2\text{SO}_4 \text{ (ammonium sulphate)} \]

- Ammonia gas can be produced when an ammonium salt is heated with an alkali e.g. \( \text{NH}_4\text{Cl} + \text{NaOH} \rightarrow \text{NaCl} + \text{H}_2\text{O} + \text{NH}_3 \)

**Calculations (separate Calculations booklet)**

- Knowing the formula for a substance, the **Formula Mass** can be calculated using relative atomic masses using values contained in the Data Booklet.
- The mass of **1 mole** of a substance is equal to the **Formula Mass** expressed in grammes - the grammie formula mass (**gfm**).
- The **mass** of any number of moles of a chemical can be calculated:

\[ \text{mass} = \text{no. of moles} \times \text{gfm} \]

- The **number of moles** in any **mass** of a chemical can be calculated:

\[ \text{no. of moles} = \frac{\text{mass}}{\text{gfm}} \]

- Calculations involving **solutions** must have volumes expressed in **litres**.

\[ \text{concentration} = \frac{\text{no. of moles}}{\text{volume}} \]

\[ \text{no. of moles} = \text{concentration} \times \text{volume} \]

- **Titration** calculations can be done in steps or by using formulae such as:

\[ C_{\text{ACID}} \times V_{\text{ACID}} \times p_{\text{ACID}} = C_{\text{ALK}} \times V_{\text{ALK}} \times p_{\text{ALK}} \quad \text{where} \ p = \text{‘power’ of the acid (H\(^+\)) or alkali (OH\(^-\))} \]

\[ \frac{C_{\text{ACID}} \times V_{\text{ACID}}}{n_{\text{ACID}}} = \frac{C_{\text{ALK}} \times V_{\text{ALK}}}{n_{\text{ALK}}} \quad \text{where} \ n = \text{number of moles in the balanced equation} \]
4.1 Acids & Bases

**Common Acids**

Acids are all around you. Most are safe to handle but some are definitely not.

**Fizzy** drinks (CO₂, carbonic acid), some fruits like lemons and oranges (citric acid), vinegar and ketchups (ethanoic acid), and even many vegetables (ascorbic acid, vitamin C) contain acids. These are examples of weak acids. A more unpleasant example is a bee sting, nettle sting or ant bite (formic acid).

Stronger acids are found in batteries (sulfuric acid) while our stomachs contain acid (hydrochloric) to break down our food into smaller molecules.

- **Acids taste sour**  
  - the latin word for sour was “acidus”

- **Acids kill cells**  
  - the most vulnerable part of you is your eyes, because they have living cells on the surface (Safety Glasses !)

  Pickling food in vinegar (ethanoic acid) is an ancient way of killing bacteria and keeping food. It made food taste sour, but many people like that, so we still pickle food.

- **Acids react with metals**  
  - iron objects, like the Forth Rail Bridge, rust much faster these days because of acid rain.

- **Acids react with carbonates**  
  - acid rain is destroying many of the marble (calcium carbonate) statues in the world

- **Acids react with alkalis**  
  - the leaves of a dock (or docken) plant can neutralise the sting from a nettle

- **Acids change the colour of indicators**  
  - e.g. they turn universal indicator from green to red

- **Acids have a pH less than 7**  
  - on a pH scale that goes from 1 to 14.
**Common Bases**

**Bases** are also all around you. Most are safe to handle but some are definitely not.

*Toothpaste* contains a base to help neutralise the *acid* on your *teeth*, produced by *bacteria*. Most *soaps* and *detergents* contain bases to help cope with *greasy* and *oily* stains. Our *liver* produces *bile* (a base) to help break down *fatty* foods. *Farmers* and *gardeners* will spread *lime* (*calcium hydroxide*) on the *soil* if it is too *acidic*.

*Wasp* stings are basic, and just as painful as any acid. Other harmful *bases* are found in *bleaches* (*ammonia*)

- **Bases feel slippery** - most *soaps* are basic
- **Bases kill cells** - the most vulnerable part of you is your *eyes*, because they have *living* cells on the *surface* (Safety Glasses !)
- **Bases react with acids** - *wasp* stings (*alkali*) should be treated with an acid like *vinegar* or *lemon* juice, while *bee* stings need *baking soda* (a base).

*Acidic* fumes (*SO$_2$* and *CO$_2$*) from *coal* burning *power stations* are passed through *lime* to be neutralised.

The *human* stomach produces *hydrochloric* acid to help the *enzyme* (*catalyst*) *pepsin* to break down *protein*. Sometimes too much *acid* is made and it begins to attack the stomach *wall* causing *pain*. All stomach remedies contain *bases* to *neutralise* the stomach acid.

- **Bases change the colour of indicators** - e.g. they turn *universal indicator* from *green* to *purple*
- **Bases have a pH greater than 7** - on a pH scale that goes from 1 to 14.
Our main source of bases are the Metal Oxides.

- lime: calcium oxide
- soda: sodium oxide
- magnesia: magnesium oxide
- pearl ash: potassium oxide

Though, we also use many Metal Carbonates.

- limestone: calcium carbonate
- marble: calcium carbonate
- baking soda: sodium carbonate
- potash: potassium carbonate

**Bases are mainly metal oxides and metal carbonates**

All of these bases are ionic compounds and, therefore, solids at room temperature. As solids, they can be directly added to an acid and will neutralise the acid.

\[
\text{base} + \text{acid} \rightarrow \text{salt} + \text{water}
\]

Sometimes, however we prefer to use solutions of soluble bases which we then can call alkalis.

Most alkalis are made by dissolving a metal oxide in water - though only those in Group 1 are very soluble. (Data Booklet)

\[
\begin{align*}
\text{Na}_2\text{O}_{(s)} + \text{H}_2\text{O}_{(l)} & \rightarrow \text{NaOH}_{(aq)} \\
\text{K}_2\text{O}_{(s)} + \text{H}_2\text{O}_{(l)} & \rightarrow \text{KOH}_{(aq)} \\
\text{CaO}_{(s)} + \text{H}_2\text{O}_{(l)} & \rightarrow \text{Ca(OH)}_2_{(aq)} \\
\text{MgO}_{(s)} + \text{H}_2\text{O}_{(l)} & \rightarrow \text{Mg(OH)}_2_{(aq)}
\end{align*}
\]

**The soluble metal oxides can form alkalis with water**
Making Acids

Probably one of the first times the word *acid* was met was in the phrase “*acid rain*”.

**Gases** like nitrogen dioxide, NO₂, (produced in petrol engines) and sulfur dioxide, SO₂, (coal burning power stations) both dissolve to produce *acid solutions*.

\[
\text{SO}_2 \ (g) \ + \ \text{H}_2 \text{O} \ (l) \ \rightarrow \ \text{H}_2\text{SO}_3 \ (aq)
\]

**Fizzy** drinks are *acidic* because carbon dioxide, CO₂, dissolves to form *carbonic* acid, H₂CO₃.

Other *non-metal oxides* like SO₃ (*sulfuric acid*, H₂SO₄) and P₂O₅ (*phosphoric acid*) behave this way, though *insoluble oxides* like carbon monoxide, CO, cannot form *acid* solutions.

---

**The soluble non-metal oxides can form acids with water**

\[
\begin{align*}
\text{SO}_2 \ (g) \ + \ \text{H}_2 \text{O} \ (l) & \rightarrow \ \text{H}_2\text{SO}_3 \ (aq) \\
\text{SO}_3 \ (g) \ + \ \text{H}_2 \text{O} \ (l) & \rightarrow \ \text{H}_2\text{SO}_4 \ (aq) \\
\text{CO}_2 \ (g) \ + \ \text{H}_2 \text{O} \ (l) & \rightarrow \ \text{H}_2\text{CO}_3 \ (aq) \\
\text{NO}_2 \ (g) \ + \ \text{H}_2 \text{O} \ (l) \ + \ \text{O}_2 \ (g) & \rightarrow \ \text{HNO}_3 \ (aq) \\
\text{P}_2\text{O}_5 \ (s) \ + \ \text{H}_2 \text{O} \ (l) & \rightarrow \ \text{H}_3\text{PO}_4 \ (aq)
\end{align*}
\]

Not all of our acids are made from oxides, however.

\[
\begin{align*}
\text{H}_2 \ (g) \ + \ \text{Cl}_2 \ (g) & \rightarrow \ \text{HCl} \ (g) \\
\text{H}_2 \ (g) \ + \ \text{S} \ (s) & \rightarrow \ \text{H}_2\text{S} \ (aq)
\end{align*}
\]

---

*Hydrochloric acid*

*Phosphorus acid*
Q1. Int2

Acids and Bases are found everywhere around us. For each of the substances listed below, decide whether they are acids (A) or bases (B).

- lemon juice (A)
- wasp stings (B)
- tomato ketchup (A)
- baking soda (B)
- toothpaste (B)
- vitamin C (B)
- Coca Cola (B)
- nettle stings (B)
- bleaches (B)
- stomach juices (B)
- detergents (B)

Q2. S

The grid shows the formulae of some oxides.

A: ZnO  B: NO₂  C: K₂O
D: CaO  E: Fe₂O₃  F: CO

a) Identify the two oxides which are covalent.

b) Identify the oxide which dissolves in water to give an alkaline solution. You may wish to use the data booklet to help you.

Q3. Int2

The grid shows the names of some compounds.

A: lead sulphate  B: sodium chloride
C: calcium hydroxide  D: potassium phosphate

a) Identify the compound which contains only two elements.

b) Identify the compound which will neutralise an acid.

Q4. SC

The grid shows the formulae of some oxides.

A: ZnO  B: NO₂  C: K₂O
D: CaO  E: Fe₂O₃  F: CO

a) Identify the two oxides which are covalent.

b) Identify the oxide which dissolves in water to give an alkaline solution. You may wish to use the data booklet to help you.

Q5. Int2

The grid shows the names of some oxides.

A: Magnesium oxide  B: Sulphur dioxide  C: Copper(II) oxide
D: Sodium oxide  E: Silicon dioxide  F: Calcium oxide
G: Iron(III) oxide  H: Potassium oxide  I: Carbon dioxide

a) Give the four boxes in the grid containing chemicals that would make an alkaline solution.

b) Give the two boxes in the grid containing chemicals that would make an acidic solution.

c) Give the three boxes in the grid containing chemicals that would make neither an acidic nor an alkaline solution.

d) Write a balanced equation for the reaction of one of the chemicals you chose in a), showing the formation of the alkali.

e) Write a balanced equation for the reaction of one of the chemicals you chose in b), showing the formation of the acid.
4.2 Acid & Base Structures

Acid Molecules

If we look carefully at the structures of the substances that dissolve in water to produce acidic solutions we can see a pattern emerge.

Firstly, they are all covalent molecules but all have a very polar bond involving a hydrogen atom.

\[
\begin{align*}
\text{hydrogen chloride} & : \quad H^+ \text{Cl}^- \\
\text{hydrogen carbonate} & : \quad H^+ \text{O}^\delta- \text{C}^\delta- \text{O}^\delta- \text{H}^+ \\
\text{hydrogen sulfate} & : \quad H^+ \text{O}^\delta- \text{S}^\delta- \text{O}^\delta- \text{O}^\delta- \text{H}^+ \\
\text{hydrogen nitrate} & : \quad \text{O}^\delta- \text{N}^\delta- \text{O}^\delta- \text{O}^\delta- \text{H}^+ 
\end{align*}
\]

Secondly, most of these molecules have a double covalent bond from the central atom to an oxygen atom next to the polar \( \text{O}^- \text{H}^+ \) bond.

\[
\begin{align*}
\text{hydrogen carbonate} & : \quad H^+ \text{O}^\delta- \text{C}^\delta- \text{O}^\delta- \text{H}^+ \\
\text{hydrogen sulfate} & : \quad H^+ \text{O}^\delta- \text{S}^\delta- \text{O}^\delta- \text{O}^\delta- \text{H}^+ \\
\end{align*}
\]

Since water molecules are also polar, there will be strong attractions set up between the water molecules and the acid molecules.

As a result, the acid molecules will be very soluble in water.

However, whilst water cannot be electrolysed at low voltages, solutions of these acid molecules can be electrolysed and always produce hydrogen gas at the negative electrode (cathode).

\[
2\text{H}^+_{(aq)} + 2e^- \rightarrow \text{H}_2(g)
\]

This suggests that the following change has taken place:

\[
\begin{align*}
\text{covalent molecule} & \rightarrow \text{ionic solution} \\
\text{acids are substances which dissolve in water to produce hydrogen ions, } H^+_{(aq)}.
\end{align*}
\]
The strong attractions between water molecules and the acid molecules do much more than simply make them soluble in water.

Water molecules are able to pull the acid molecules apart.

In the process, the shared electrons are completely transferred to the atom with the stronger pull - in this case the chlorine atom.

The acid molecule is ionised - turned into ions dissolved in water.

Acids are substances which dissolve in water to produce hydrogen ions, $H^+_{(aq)}$. 

<table>
<thead>
<tr>
<th>Chemical Formula</th>
<th>Molecular Structure</th>
<th>Notes</th>
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<tbody>
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<td>$H\delta^+_1Cl\delta^-$</td>
<td><img src="" alt="hydrogen chloride" /></td>
<td>hydrogen chloride</td>
</tr>
<tr>
<td>$(H^+)_2CO_3^{2-}$</td>
<td><img src="" alt="hydrogen carbonate" /></td>
<td>carbonic acid</td>
</tr>
<tr>
<td>$(H^+)_2SO_4^{2-}$</td>
<td><img src="" alt="hydrogen sulfate" /></td>
<td>sulfuric acid</td>
</tr>
<tr>
<td>$H^+NO_3^{-}$</td>
<td><img src="" alt="hydrogen nitrate" /></td>
<td>nitric acid</td>
</tr>
</tbody>
</table>
Ammonia

Ammonia, is the 'only' base that starts off as a covalent molecule but when it dissolves in water the solution conducts showing that ions are formed. The pH of the solution is > 7 showing that it has formed an alkaline solution.

The two polar molecules attract each other strongly enough to pull a hydrogen ion (H+) off the water molecule.

The resulting OH⁻ ion makes the solution alkali.

\[
\text{NH}_3(\text{g}) + \text{H}_2\text{O}(\text{l}) \rightarrow \text{NH}_4^+\text{OH}^- \quad \text{(aq)}
\]

A convenient way to make ammonia in the laboratory is to heat an ammonium compound with a strong base or alkali. (The reverse of the reaction above)

Properties of Ammonia

Ammonia is a colourless gas at 20 °C
Ammonia is the only alkali gas
Ammonia has a strong pungent smell
Ammonia is less dense than air
Ammonia dissolves in water to form an alkaline solution which will conduct electricity

As the ammonia gas cools and contracts, the water is slowly sucked up the tube. As the first drop appears at the end of the tube ALL the ammonia gas dissolves in this single drop of water.

As a result, a partial vacuum exists in the flask resulting in more water being sucked in very quickly to replace the ammonia gas molecules.

A 'fountain' is the result, showing that ammonia is extremely soluble in water. This is known as the Fountain Experiment.
Water Ions

More surprisingly, perhaps, is the fact that *attractions* between *water molecules* can also result in this *covalent molecule* being pulled apart to form *ions*.

The two *polar molecules* attract each other *strongly* enough to pull a *hydrogen ion* (H\(^+\)) off one of the *water* molecules.

However an OH\(^-\) ion is also formed so water remains overall *neutral*, pH = 7.

\[
\begin{align*}
\text{Dissociation} & : \ H_2O \ (l) \ \rightarrow \ H^+ \ (aq) + OH^- \ (aq) \\
\text{Almost immediately}, & \ \text{the ions will attract each other and the molecule will reform.} \\
\text{Water is a mixture of molecules and ions, constantly breaking up and reforming.} \\
\text{The vast majority of water is molecular but there are always enough ions to make water a poor conductor.}
\end{align*}
\]

*Acids are substances which dissolve in water to increase the concentration of hydrogen ions, H\(^+(aq)\) - H\(^+\) concentration > 10\(^{-7}\)*

*Bases are substances which dissolve in water to increase the concentration of hydroxide ions, OH\(^-(aq)\) - OH\(^-\) concentration > 10\(^{-7}\)*


**pH Numbers**

Neutral solutions have a pH which equals 7, pH = 7.

Neutral solutions have equal amounts of $H^+$ ions and $OH^-$ ions

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<tr>
<th>$H^+$</th>
<th>$10^0$</th>
<th>$10^{-1}$</th>
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<td>$OH^-$</td>
<td>$10^{-14}$</td>
<td>$10^{-15}$</td>
<td>$10^{-16}$</td>
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<td>$10^{-27}$</td>
<td>$10^{-28}$</td>
<td>$10^{-29}$</td>
</tr>
</tbody>
</table>

**Acid solutions** have a pH which is less than 7, pH < 7.

**Alkali solutions** have a pH which is more than 7, pH > 7.

**Acid solutions** have more $H^+$ ions than $OH^-$ ions, $H^+ > OH^-$

**Alkali solutions** have more $H^+$ ions than $OH^-$ ions, $H^+ > OH^-$

If you dilute an acid by adding more water, the pH will increase towards pH = 7.

If you dilute an alkali by adding more water, the pH will decrease towards pH = 7.

It would be tempting to assume that an acid of pH = 1, eg stomach acid, was three times as concentrated as an acid of pH = 3, e.g. lemon juice.

In fact, it is $10 \times 10$ i.e. 100 times as concentrated.

For each change in pH, a solution will become ten times more concentrated or ten times less concentrated.

pH paper, universal indicator and other coloured substances, such as red and blue litmus paper and even the juice from red cabbage, can be used to measure the pH of solutions.

Conductivity devices, (H⁺ ions are good conductors), such as pH meters and pH probes can also be used.
Q1.

a) What two features are found in most acid molecules?


b) Which of the following is an acid molecule?

A

B

C

D


Q2.

The grid shows some statements which can be applied to different solutions.

A  It has a pH less than 7.
B  It conducts electricity.
C  It contains less OH\(^{-}\) ions than pure water.
D  It does not neutralise dilute hydrochloric acid.
E  When diluted the concentration of OH\(^{-}\) ions decreases.

Identify the two statements which are correct for an alkaline solution.


Q3.

What is the most likely pH value that would be obtained when zinc oxide is added to water?

(You may wish to use page 5 of the data booklet to help you.)

A  5
B  7
C  9
D  11

Q4.

For each of the acids below

a) Write the formula for the molecule

b) Write the formulae for the two ions formed when it dissolves (dissociates) in water?

\[
\text{H-O-S-O-H} \\
\text{H-O-S-O-H} \\
\text{H-O-S-O-H} \\
\text{H-O-S-O-H}
\]

\[
\text{O=O-N-O-H} \\
\text{H-N-C-O-H} \\
\text{H-N-C-O-H} \\
\text{H-N-C-O-H}
\]

\[
\text{Cl-O-C-O-H} \\
\text{H-C-O-C-O-H} \\
\text{H-C-O-C-O-H} \\
\text{H-C-O-C-O-H}
\]

Q5.

A solution of 0·1 mol/l hydrochloric acid has a pH of 1.

a) What colour would universal indicator turn when added to a solution of hydrochloric acid?

b) Starting at pH 1, draw a line to show how the pH of this acid changes when diluted with water.

c) Calculate the number of moles of hydrochloric acid in 50 cm\(^3\) of 0·1 mol/l hydrochloric acid solution.

d) Magnesium carbonate can be used to neutralise acid:

\[
\text{MgCO}_3(\text{s}) + 2 \text{HCl}_{(\text{aq})} \rightarrow \text{MgCl}_2(\text{aq}) + \text{H}_2\text{O}_{(l)} + \text{CO}_2(\text{g})
\]

i) Calculate the number of moles of MgCO\(_3\) needed to neutralise 50 cm\(^3\) of 0·1 mol/l HCl\(_{\text{aq}}\).

ii) Calculate the mass of MgCO\(_3\) needed to neutralise 50 cm\(^3\) of 0·1 mol/l HCl\(_{\text{aq}}\).
4.3 Reactions of Acids

With Metals

As you have probably learnt in earlier courses, reactive metals that are above hydrogen in the Reactivity Series are able to react with acids.

The gas produced burns with a squeaky-pop. This shows that the gas is again hydrogen.

Reactive metals are able to force hydrogen ions to change back to hydrogen atoms. This allows the metal to take the place of the hydrogen and form a new substance called a Salt.

The sodium ion takes the place of the hydrogen ion to form the salt called sodium chloride.

Each acid has its own salts:-

- hydrochloric acid, HCl $\rightarrow$ chlorides e.g. sodium chloride, NaCl
- sulfuric acid, H$_2$SO$_4$ $\rightarrow$ sulfates e.g. copper sulfate, CuSO$_4$
- nitric acid, HNO$_3$ $\rightarrow$ nitrates e.g. potassium nitrate, KNO$_3$

<table>
<thead>
<tr>
<th>Acid</th>
<th>+</th>
<th>metal</th>
<th>$\rightarrow$</th>
<th>salt</th>
<th>+</th>
<th>hydrogen</th>
</tr>
</thead>
<tbody>
<tr>
<td>e.g. magnesium</td>
<td>+</td>
<td>sulfuric acid</td>
<td>$\rightarrow$ magnesium</td>
<td>+</td>
<td>hydrogen sulfate</td>
<td></td>
</tr>
<tr>
<td>Mg$_{(s)}$</td>
<td>+</td>
<td>H$_2$SO$_4^{(aq)}$</td>
<td>$\rightarrow$ MgSO$_4^{(aq)}$</td>
<td>+</td>
<td>H$_2^{(g)}$</td>
<td></td>
</tr>
</tbody>
</table>

We will learn more about this reaction if we firstly write an ionic equation and then remove spectator ions.

e.g. Mg$_{(s)}$ + (H$^+$)$_2$SO$_4^{2-}$$_{(aq)}$ $\rightarrow$ Mg$^{2+}$SO$_4^{2-}$$_{(aq)}$ + H$_2^{(g)}$

Mg$_{(s)}$ + 2 H$^+$(aq) $\rightarrow$ Mg$^{2+}$(aq) + H$_2^{(g)}$
With Metal Oxides

In **metal oxides** the **metal** has already formed an **ion** and will not **react** any more.

The **oxide ion**, $\text{O}^{2-}$, reacts with the **hydrogen ions** in the **acid** to form **water**, $\text{H}_2\text{O}$.

This leaves the **metal ion** to take the **hydrogen ions** place, so again a **salt** will be produced.

<table>
<thead>
<tr>
<th>Acid</th>
<th>+</th>
<th>metal oxide</th>
<th>→</th>
<th>salt</th>
<th>+</th>
<th>water</th>
</tr>
</thead>
<tbody>
<tr>
<td>e.g. iron (III) oxide</td>
<td>+</td>
<td>nitric acid</td>
<td>→</td>
<td>iron(III) nitrate</td>
<td></td>
<td></td>
</tr>
<tr>
<td>balanced</td>
<td>Fe$_2$O$_3$ (s) + 6 HNO$_3$ (aq) → 2 Fe(NO$_3$)$_3$ (aq) + 3 H$_2$O (l)</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>ionic</td>
<td>(Fe$^{3+}$)$_2$(O$^{2-}$)$_3$ (s) + 6 H$^+$NO$_3^-$ (aq) → 2 Fe$^{3+}$(NO$_3^-$)$_3$ (aq) + 3 H$_2$O (l)</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>without spectator ions</td>
<td>3 O$^{2-}$ (s) + 6 H$^+$ (aq) → 3 H$_2$O (l)</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

With Metal Carbonates

In **metal carbonates** the **metal** has already formed an **ion** and will not **react** any more.

The **carbonate ion**, $\text{CO}_3^{2-}$, reacts with the **hydrogen ions** in the **acid** to form **water**, $\text{H}_2\text{O}$ and **carbon dioxide** gas, $\text{CO}_2$.

This leaves the **metal ion** to take the hydrogens place, so again a **salt** will be produced.

<table>
<thead>
<tr>
<th>Acid</th>
<th>+</th>
<th>metal carbonate</th>
<th>→</th>
<th>salt</th>
<th>+</th>
<th>water</th>
<th>+</th>
<th>carbon dioxide</th>
</tr>
</thead>
<tbody>
<tr>
<td>e.g. calcium carbonate</td>
<td>+</td>
<td>hydrochloric acid</td>
<td>→</td>
<td>calcium chloride</td>
<td>+</td>
<td>carbon dioxide</td>
<td></td>
<td></td>
</tr>
<tr>
<td>balanced</td>
<td>CaCO$_3$ (s) + 2 HCl (aq) → CaCl$_2$ (aq) + H$_2$O (l) + CO$_2$ (g)</td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

KHS Oct 2013
Acids & Bases

**Topic 4**

### National 5

**With Alkalis**

*Alkalis* are *solutions* which contain *hydroxide ions*.

Most *alkalis* are made by *dissolving* a *metal oxide* in *water* - though only those in *Group 1* are very *soluble*.

\[
\begin{align*}
\text{Na}_2\text{O} \; (s) & \quad + \quad \text{H}_2\text{O} \; (l) \\
& \quad \rightarrow \\
\text{NaOH} \; (aq) & \\
\text{soda} & \\
\text{K}_2\text{O} \; (s) & \quad + \quad \text{H}_2\text{O} \; (l) \\
& \quad \rightarrow \\
\text{KOH} \; (aq) & \\
\text{pearl ash} & \\
\text{CaO} \; (s) & \quad + \quad \text{H}_2\text{O} \; (l) \\
& \quad \rightarrow \\
\text{Ca(O}_2\text{)} \; (aq) & \\
\text{lime} & \\
\text{MgO} \; (s) & \quad + \quad \text{H}_2\text{O} \; (l) \\
& \quad \rightarrow \\
\text{Mg(OH)} \; (aq) & \\
\text{magnesia} & \\
\end{align*}
\]

It is the *hydroxide ion* which will react with the *hydrogen ion* in the *acid*, and *water* is the product.

### Acid + alkali → salt + water

<table>
<thead>
<tr>
<th>Acid</th>
<th>alkali</th>
<th>salt</th>
<th>water</th>
</tr>
</thead>
<tbody>
<tr>
<td>sodium hydroxide</td>
<td>+</td>
<td>sulfuric acid</td>
<td>→ sodium sulfate</td>
</tr>
<tr>
<td>2 NaOH ( (aq) ) + H(_2)SO(_4) ( (aq) )</td>
<td>( \rightarrow )</td>
<td>Na(_2) SO(_4) ( (aq) ) + 2 H(_2)O ( (l) )</td>
<td></td>
</tr>
<tr>
<td>ionic</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>2 Na(^+)OH(^-) ( (aq) ) + (H(^+))(_2) SO(_4) (^{2-}) ( (aq) )</td>
<td>( \rightarrow )</td>
<td>(Na(^+))(_2) SO(_4) (^{2-}) ( (aq) ) + 2 H(_2)O ( (l) )</td>
<td></td>
</tr>
</tbody>
</table>

Once again, the *metal ion* to take the hydrogens place, so again a *salt* will be produced. The *metal ions* are *spectator ions* as are the *sulfate ions* in the example above.

**without spectator ions**

\[
2\text{OH}^- \quad (aq) \quad + 
2\text{H}^+ \quad (aq) 
\rightarrow 
2\text{H}_2\text{O} \quad (l)
\]
Salt Preparation

Because of the large number of, in particular, metal oxides and carbonates that it is possible to react easily with a number of acids, a whole range of ‘new substances’ can be made using acid reactions. If we include, *precipitation* reactions (met earlier in the course) there are very few compounds that cannot be made quickly and easily.

This is basically a *Problem Solving* activity that will test your knowledge of *acid reactions*, your use of *solubility tables*, your appreciation of *practical considerations* (such as ensuring complete reaction) and your knowledge of *separation techniques*.

There are 3 parts to salt preparation  

1. **Choice of Reaction** 
2. **Reaction Method** 
3. **Separation of Salt produced**

### 1) Possible Reactions

- **a)** Acid + (solid) Metal
- **b)** Acid + (solid) Oxide, Hydroxide or Carbonate
- **c)** Acid + Alkali (solution)
- **d)** Precipitation (solutions, one of which may be acid)

### 2) Reaction Methods & 3) Separation of Salt

**Solid Metals, Oxides, Hydroxides & Carbonates**

The solid is added spatula by spatula, stirring all the time, until there is an obvious layer of unreacted (excess) solid lying at the bottom. The acid may need to be heated to speed up the reaction.

**Alkali solutions**

The acid will need to be added *slowly* and carefully (eventually drop by drop) until the indicator just changes colour.

Then, to another flask, the exact same volumes will be reacted but **without** the indicator present.

The *excess* solid must now be separated from the *salt* solution by filtering. The solid trapped in the filter paper can be discarded.

The *salt* solution can now be heated until all the water has evaporated away leaving solid *salt* powder. If preferred, the solution can be left to evaporate *slowly* in which case *salt crystals* will form.
Precipitation reactions

Using the Data Book, two suitable solutions will need to be made up.
Each solution will provide one half of the salt to be made.
Once made, the two solutions are simply mixed together.

Sometimes a choice of methods is available. To help choose the ‘best’ method the following summary may be useful.

<table>
<thead>
<tr>
<th>Reaction</th>
<th>'Advantages'</th>
<th>'Disadvantages'</th>
<th>'Suitability'</th>
</tr>
</thead>
<tbody>
<tr>
<td><strong>solid metals</strong></td>
<td>Easy to ensure 'complete' reaction - excess metal left over at end.</td>
<td>Must filter excess metal. Not all metals reactive enough to react with acid.</td>
<td>Not suitable if salt is insoluble - too difficult to separate from excess solid metal.</td>
</tr>
<tr>
<td><strong>solid oxides, hydroxides, carbonates</strong></td>
<td>Easy to ensure 'complete' reaction - excess solid left over at end.</td>
<td>Must filter excess solids. Often need to heat oxides and hydroxides.</td>
<td>Not suitable if salt is insoluble - too difficult to separate from excess solid.</td>
</tr>
<tr>
<td><strong>alkali solutions</strong></td>
<td>Reaction immediate. No need to filter excess solids.</td>
<td>Difficult to ensure 'exact' neutralisation. Technique may take a very long time.</td>
<td>Very limited choice of alkalis - so limited number of salts can be prepared by this method.</td>
</tr>
<tr>
<td><strong>precipitation from solutions</strong></td>
<td>Reaction extremely quick. None really.</td>
<td></td>
<td>Limited to insoluble salts only.</td>
</tr>
</tbody>
</table>

Examples

To prepare **copper sulfate**

1. The Data Book will tell you that copper sulfate is soluble, so precipitation is out.
2. Acid to use: **sulfuric acid**
3. Copper metal - no reaction with acid.
4. Copper oxide/hydroxide is insoluble so there is no alkali solution, so titration is out.
5. Best method would be to add solid copper oxide/hydroxide/carbonate to sulfuric acid.

To prepare **zinc chloride**

1. The Data Book will tell you that zinc chloride is soluble, so precipitation is out.
2. Acid to use: **hydrochloric acid**
3. Zinc reacts slowly with acid.
4. Zinc oxide/hydroxide is insoluble so there is no alkali solution, so titration is out.
5. Best method would be to add solid zinc or zinc oxide/hydroxide/carbonate to hydrochloric acid.

To prepare **silver chloride**

1. The Data Book will tell you that silver chloride is insoluble, so precipitation is the best method.
2. Use the Data Book to find a soluble silver compound eg **silver nitrate**.
Use the Data Book to find a soluble chloride compound eg **sodium chloride**.
Q1. SC
The grid shows the names of some metals.

Identify the metal that does not react with dilute acid.
You may wish to use page 7 of the data booklet to help you.

Q2. SC
*Copy and complete* the following equations

- lithium + hydrochloric → + water
  - hydroxide + acid
- aluminium + nitric → + water
  - oxide + acid
- strontium + nitric → + water
  - carbonate + acid
- zinc + sulphuric → +

Q3. SC
*Copy, complete and balance* the following equations

- Ca(OH)₂ + HCl → + H₂O
- CuO + → Cu(NO₃)₂ +
  + H₂SO₄ → K₂SO₄ + + CO₂
- Li + + LiCl +

Q4. SC
Lead(II) nitrate solution reacts with potassium iodide solution to give a yellow solid.

\[ \text{Pb}^{2+} (aq) + 2\text{NO}_3^- (aq) + 2\text{K}^+ (aq) + 2\text{I}^- (aq) \rightarrow \text{PbI}_2 (s) + 2\text{K}^+ (aq) + 2\text{NO}_3^- (aq) \]

Identify the *two* spectator ions in the reaction.

Q5. SC
The grid shows the names of some soluble compounds.

\[ \begin{array}{ccc}
\text{A} & \text{B} & \text{C} \\
\text{sodium iodide} & \text{potassium chloride} & \text{lithium chloride} \\
\text{D} & \text{E} & \text{F} \\
\text{barium bromide} & \text{sodium hydroxide} & \text{potassium sulphate} \\
\end{array} \]

(a) Identify the base.
(b) Identify the two compounds whose solutions would form a precipitate when mixed.
You may wish to use the data booklet to help you.

Q6. Int2
*Copy and complete* the following equations, clearly showing which of the products is the *precipitate*.

- \[ \text{BaCl}_2 (aq) + \text{K}_2\text{SO}_4 (aq) \rightarrow + \]
- \[ \text{CuSO}_4 (aq) + \text{Na}_2\text{CO}_3 (aq) \rightarrow + \]
- \[ \text{AgNO}_3 (aq) + \text{KCl} (aq) \rightarrow + \]

Q7. Int2
Reactions can be represented using ionic equations. Which ionic equation shows a neutralisation reaction?

\[ \begin{array}{cccc}
\text{A} & 2\text{H}_2\text{O} (l) + \text{O}_2 (g) + 4e^- & \rightarrow & 4\text{OH}^- (aq) \\
\text{B} & \text{H}^+ (aq) + \text{OH}^- (aq) & \rightarrow & \text{H}_2\text{O} (l) \\
\text{C} & \text{SO}_2 (g) + \text{H}_2\text{O} (l) & \rightarrow & 2\text{H}^+ (aq) + \text{SO}_3^{2-} (aq) \\
\text{D} & \text{NH}_4^+ (s) + \text{OH}^- (s) & \rightarrow & \text{NH}_3 (g) + \text{NH}_4^+ (s) \\
\end{array} \]

Q8. SC
A solution of sulphuric acid can be used to neutralise a solution of sodium hydroxide.

(a) What is the pH of the solution when it is exactly neutral.
(b) What is the name of the salt formed in the neutralisation reaction?
(c) Balance the following equation for the reaction.
\[ (\text{H}_2\text{SO}_4)^{2-} + \text{NaOH} \rightarrow (\text{Na}^+)_2\text{SO}_4^{2-} + \text{H}_2\text{O} \]
(d) Rewrite the equation, omitting the spectator ions.
4.4 Quantitative Analysis

Titrations

It is quite common for a Chemist to be asked to measure how much acid or alkali is present in a solution - for example, how much acid is present in lemonade.

A technique called titration is used, and you may be expected to demonstrate your ability to carry out a titration.

A carefully measured volume of the lemonade would be placed in the flask. An indicator that will change colour is also added.

An alkali whose concentration is accurately known would be placed in a burette, and then added to the flask.

Eventually the alkali would be added drop by drop until the indicator changes colour to show that the acid in the lemonade has been neutralised.

Before the alkali can be used to determine how much acid is present in the lemonade, it must be standardised - titrated against a Standard Solution of a suitable acid, such as potassium hydrogen phthalate (KHP).

An unusual acid like this is chosen because it is extremely stable, very soluble and can be made to a very high level of purity.

An analytical balance is used to very accurately weigh out a calculated mass of the chemical.

From this mass, the number of moles of chemical can be calculated and used to calculate the concentration of the solution made.

Molecular Formula:

\[ \text{H}_2\text{C}_6\text{O}_4\text{K}^+ \]

Formula Mass:

Mass of one mole:
Technique - Standardisation of alkali (NaOH)

1. **Filling the burette**

   - Clamp gently!

2. Collect a beakerful of the alkali you are going to use and another empty waste beaker.

3. Pour just a little of the alkali into the burette to rinse it. Pour it out into your waste beaker.

4. With the valve closed, fill up the burette to just above the zero line. Then remove the funnel.

5. Open the valve slightly and let the alkali drip into your waste beaker until the bottom of the curved surface is on or below the zero line.

6. Read your starting volume. Read with your eye level with the curved surface. Make a note of this reading. The burette is now ready for use.

---

**Mass of KHP = 10.00 g**

\[
\begin{align*}
204.22 \text{ g} & \quad \rightarrow \quad 1 \text{ mole} \\
10.00 \text{ g} & \quad \rightarrow \quad \frac{1 \times 10.00}{204.22} \\
& = 0.0490 \text{ moles}
\end{align*}
\]

**Vol of flask = 500 ml = 0.5 l**

\[
C = \frac{n}{V} = \frac{0.0490}{0.5} = 0.098 \text{ mol l}^{-1}
\]
2. **Using the pipette**

The pipette is used to accurately measure out the same **volume** of KHP every time.

1. Collect a **beakerful** of the KHP you are going to use, and a conical **flask**.
2. Use the **filler** to suck the KHP above the fill mark.
3. Holding the pipette above the beaker, slowly let the KHP drip out until the bottom of the curved surface is on the fill mark.
4. Carefully transfer the KHP in the pipette into your flask.

A tiny amount will remain inside the tip. This is supposed to happen.

Dip the tip of the pipette into the KHP and some more will come out. Any still left in the pipette is allowed for.

5. Add a few drops of **indicator**.

3. **Doing the Titration**

The aim is to find out exactly what **volume** of alkali is needed to **neutralise** the **known volume** of KHP

1. Put a piece of **white paper** under the flask. It will help you to see the **colour** of indicator.
2. Start by adding the alkali 5 ml at a time. You should see the **indicator colour** change but then return quickly.
3. If the **colour** takes longer to return, add less alkali next time. Ideally you should add one drop of alkali and see the indicator change permanently.
4. Write down the final burette reading to at least the nearest 0.1 ml.

Remember that a burette reads downwards.

The burette opposite is reading 17.8 ml (and not 18.2 ml).

5. You will now need to repeat your titration with a freshly pipetted sample of KHP.

Knowing the answer from the first attempt should allow you to quickly add enough alkali to nearly neutralise the potassium hydrogen phthalate. Then you can add more alkali drop by drop to get an accurate result.

Results

6. You should record your results in a table similar to the one on the right.

<table>
<thead>
<tr>
<th>Attempt number</th>
<th>Starting volume (ml)</th>
<th>Final volume (ml)</th>
<th>Volume added (ml)</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>0.1</td>
<td>22.0</td>
<td>21.9</td>
</tr>
<tr>
<td>2</td>
<td>22.0</td>
<td>43.5</td>
<td>21.5</td>
</tr>
<tr>
<td>3</td>
<td>0.2</td>
<td>21.9</td>
<td>21.7</td>
</tr>
<tr>
<td>4</td>
<td>21.9</td>
<td>42.6</td>
<td>21.7</td>
</tr>
</tbody>
</table>

Your first attempt is often a ‘rough’ titration as you will often add too much alkali at a time.

Later attempts should produce results to the nearest drop.

You have to continue repeating the titration until you get two results at least within 0.1 ml of each other.

7. You should finish off by quoting your conclusion in terms of:-

“It takes 21.7 ml of NaOH to neutralise 20 ml of KHP”.

If you have two, or more, answers close to each other then it may be better to use their average as your final answer.

To get similar results each time, (the aim of this technique), you will need to work hard to ensure that you pipette exactly the same volume of potassium hydrogen phthalate each time.
Evaluation

As part of the Assessment Task 1, you will be asked to evaluate experimental procedures, mentioning at least one of the below.

- effectiveness of procedure
  
  the effectiveness of a titration, for example, can often be determined by:
  
  how often you had to repeat the titration? how was the colour change (1 drop)?
  
  how close to each other were your volumes?

- control of variables
  
  how well did you control variables such as:
  
  concentration of NaOH, volume of KHP, amount of indicator added etc?

- limitations of equipment
  
  any issues with equipment such as:
  
  electronic balance (2 or 3 decimal place?), volumetric flasks (A or B grade?),
- **sources of uncertainty**

these are often 'built-in errors' such as:
- **weighings** \((\pm \ ? \ g)\), **volumetric flasks** \((\pm \ ? \ cm^3)\), **pipettes** \((\pm \ ? \ cm^3)\), **burettes** \((\pm \ ? \ cm^3)\), ** burette readings** \((\pm \ ? \ cm^3)\) etc - we often express these as % errors.

- **possible improvements**

these will often emerge from the previous categories:
Q1.

The table below shows the colours of various indicators at different pH values.

<table>
<thead>
<tr>
<th>indicator</th>
<th>pH 1</th>
<th>colour 1</th>
<th>pH 2</th>
<th>colour 2</th>
</tr>
</thead>
<tbody>
<tr>
<td>bromophenol blue</td>
<td>3</td>
<td>yellow</td>
<td>4.5</td>
<td>blue</td>
</tr>
<tr>
<td>phenolphthalein</td>
<td>8</td>
<td>colourless</td>
<td>10</td>
<td>pink</td>
</tr>
<tr>
<td>methyl orange</td>
<td>3</td>
<td>red</td>
<td>4.5</td>
<td>yellow</td>
</tr>
<tr>
<td>thymol blue</td>
<td>6</td>
<td>yellow</td>
<td>7.5</td>
<td>blue</td>
</tr>
</tbody>
</table>

The table below shows the pH of some solutions.

<table>
<thead>
<tr>
<th>solution</th>
<th>pH</th>
</tr>
</thead>
<tbody>
<tr>
<td>0.1 M hydrochloric acid</td>
<td>1.0</td>
</tr>
<tr>
<td>0.1 M ethanoic acid</td>
<td>5.0</td>
</tr>
<tr>
<td>0.1 M ammonia</td>
<td>10.0</td>
</tr>
<tr>
<td>0.1 M sodium hydroxide</td>
<td>12.5</td>
</tr>
</tbody>
</table>

a) Complete the table below to show the colours of the indicators in the solutions.

<table>
<thead>
<tr>
<th>indicator</th>
<th>solution</th>
<th>colour</th>
</tr>
</thead>
<tbody>
<tr>
<td>bromophenol blue</td>
<td>0.1 M hydrochloric acid</td>
<td></td>
</tr>
<tr>
<td>phenolphthalein</td>
<td>0.1 M ethanoic acid</td>
<td></td>
</tr>
<tr>
<td>methyl orange</td>
<td>0.1 M ammonia</td>
<td></td>
</tr>
<tr>
<td>thymol blue</td>
<td>0.1 M sodium hydroxide</td>
<td></td>
</tr>
</tbody>
</table>

b) Name one indicator which turns the same colour in both ethanoic acid and sodium hydroxide.

______________

c) Which two indicators turn the same colour in hydrochloric acid

______________

Q2.

One of the solids often used in Antacid Tablets to treat indigestion is magnesium hydroxide.

A pupil decided to find out how much of the solid would be needed to neutralise some acid.

a) Complete the equation for the reaction of magnesium hydroxide with hydrochloric acid

\[ \text{Mg(OH)}_2 (aq) + 2\text{HCl} (aq) \rightarrow + \text{H}_2\text{O} (l) \]

b) Calculate the number of moles of HCl present.

c) Calculate the number of moles of Mg(OH)₂ needed.

d) Calculate the mass of Mg(OH)₂ needed.
Q1.

Below is information about six chemicals.

<table>
<thead>
<tr>
<th>chemical</th>
<th>state at 20 °C</th>
<th>pH in water</th>
<th>reaction with water</th>
</tr>
</thead>
<tbody>
<tr>
<td>A</td>
<td>gas</td>
<td>1</td>
<td>none</td>
</tr>
<tr>
<td>B</td>
<td>liquid</td>
<td>7</td>
<td>none</td>
</tr>
<tr>
<td>C</td>
<td>solid</td>
<td>4</td>
<td>none</td>
</tr>
<tr>
<td>D</td>
<td>solid</td>
<td>8</td>
<td>forms salt, carbon dioxide and water</td>
</tr>
<tr>
<td>E</td>
<td>solid</td>
<td>14</td>
<td>forms a salt and water</td>
</tr>
<tr>
<td>F</td>
<td>solid</td>
<td>no reaction</td>
<td>fizzes</td>
</tr>
</tbody>
</table>

Use the table to write the letter of the chemical substance which:

a) forms the most strongly acidic solution ______

b) forms a neutral solution ______

c) is a metal ______

d) forms a solution which turns pH paper orange ______

e) is a carbonate ______

f) is water ______

g) is sulphur dioxide ______

Q2. SC

Equations are used to represent chemical reactions.

A) \( Zn(s) \rightarrow Zn^{2+}(aq) + 2e^- \)

B) \( C_2H_5OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 3H_2O(l) \)

C) \( SO_2(g) + H_2O(l) \rightarrow 2H^+(aq) + SO_4^{2-}(aq) \)

D) \( H^+(aq) + OH^-(aq) \rightarrow H_2O(l) \)

E) \( SO_3^{2-}(aq) + 2H^+(aq) + 2e^- \rightarrow SO_4^{2-}(aq) + H_2O(l) \)

a) Identify the equation which represents the formation of acid rain.

b) Identify the equation which involves sulphuric acid.

Q3. Int2

Reactions can be represented using ionic equations. Which ionic equation shows a neutralisation reaction?

A) \( 2H_2O(l) + O_2(g) + 4e^- \rightarrow 4OH^- \)

B) \( H^+ + OH^- \rightarrow H_2O(l) \)

C) \( SO_2(g) + H_2O(l) \rightarrow 2H^+ + SO_4^{2-} \)

D) \( NH_4^+ + OH^- \rightarrow NH_3(g) + NH_4^+ \)

Q4. SC

The grid shows some ions.

A) \( Al^{3+} \)

B) \( Cl^- \)

C) \( Li^+ \)

D) \( H^+ \)

E) \( Br^- \)

F) \( OH^- \)

a) Identify the two ions which combine to form an insoluble compound.

b) You may wish to use the data booklet to help you.

Q5. SC

Lead(II) nitrate solution reacts with potassium iodide solution to give a yellow solid.

\[ Pb^{2+} + 2NO_3^- + 2K^+ + 2I^- \rightarrow PbI_2(s) + 2K^+ + 2NO_3^- \]

Identify the two spectator ions in the reaction.

KHS Oct 2013
CONSOLIDATION QUESTIONS  

Q1 Both ammonia molecules and hydrogen chloride molecules are described as being polar.

   a) What is meant by the word polar, as used in this context.

   ________________________________________________________________
   ________________________________________________________________

   b) Complete the formula for hydrogen chloride to show its polar characteristics.

   \[ \text{H} \quad \text{–} \quad \text{Cl} \]

   c) Ammonia gas \( \text{NH}_3(\text{g}) \), can be dissolved in water to form concentrated ammonia solution. Hydrogen chloride gas \( \text{HCl}(\text{g}) \), can be dissolved in water to form concentrated hydrochloric acid.

   If both bottles are placed next to each other in a fume cupboard and the stoppers removed, both liquids evaporate and a white cloud is formed where the two gases meet.

   i) State the colour of the pH paper at X and Y.

   pH paper X __________________ pH paper Y __________________

   ii) The white cloud appears because the gases react to form a salt. Name the salt formed.

   ________________________________________________________________
CONSIDERATION QUESTIONS

Q1 A student investigated the reaction between dilute sulphuric acid and sodium carbonate.

His experiment involved determining the concentration of sodium carbonate solution by titration.

The results showed that 20 cm$^3$ of sulphuric acid was required to neutralise the sodium carbonate solution.

a) Calculate the number of moles of sulphuric acid in this volume.

\[ \text{mol} \]

b) One mole of sulphuric acid reacts with one mole of sodium carbonate.

Using your answer from part a), calculate the concentration, in mol/l, of the sodium carbonate solution.

\[ \text{mol/l} \]

c) Name the salt produced when dilute sulphuric acid reacts with sodium carbonate.

_________________________________________________________________________